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Q1. If the door is left open, the kinetic energy of the hot air outside will be transferred more easily to the cool air inside than if the door is closed. Moving air transfers thermal energy more easily than the particles of the door.

Q2. When you touch an object that has a lower temperature than your skin, the thermal energy from your skin is transferred to the colder object. The temperature of the cooler object will increase, and the temperature of your skin will decrease. Eventually, the temperature of both objects will reach the same temperature.

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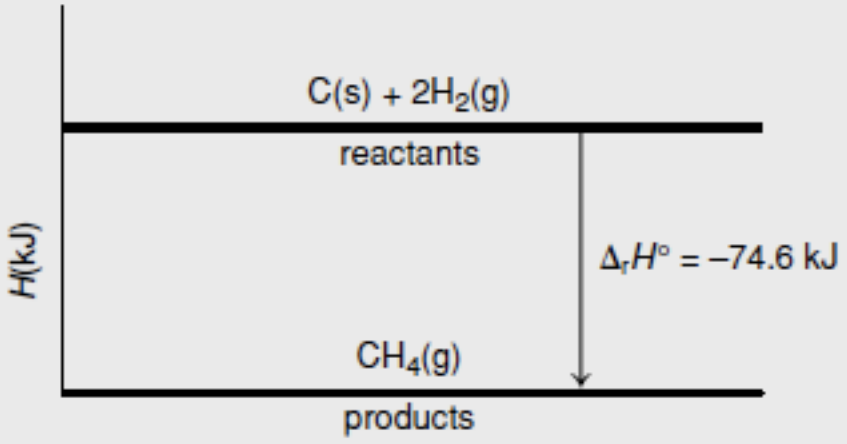
Q3. (a) $C(s) + 2H_2(g) \rightarrow CH_4(g) + 74.6\ kJ$

(b) $C(s) + 2H_2(g) \rightarrow CH_4(g)\ \Delta_r H^\circ = -74.6\ kJ$

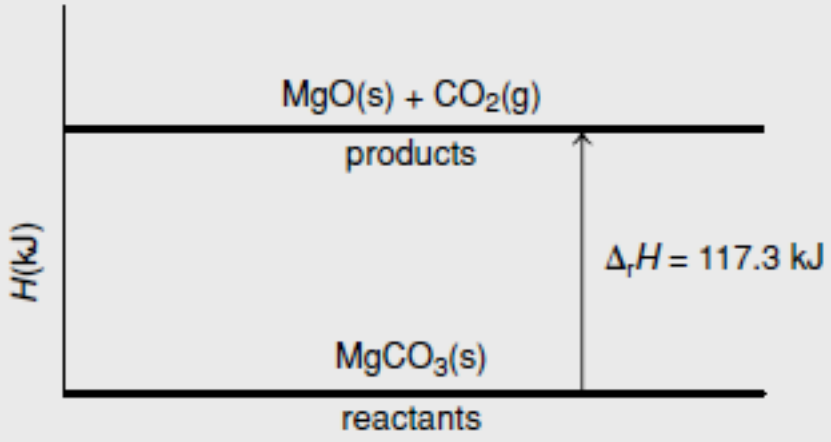
(c)



Q4. (a) $MgCO_3(s) + 117.3\ kJ \rightarrow MgO(s) + CO_2(g)$

(b) $MgCO_3(s) \rightarrow MgO(s) + CO_2(g)\ \Delta_r H = 117.3\ kJ$

(c)



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Q5. (a) $C_5H_{12}(\ell) + 8O_2(g) \rightarrow 5CO_2(g) + 6H_2O(g)$

$\Delta_c H = -3244.8$ kJ/mol

(b) $C_4H_{10}(g) + \frac{13}{2}O_2(g) \rightarrow 4CO_2(g) + 5H_2O(g)$

$\Delta_c H = -2657.3$ kJ/mol

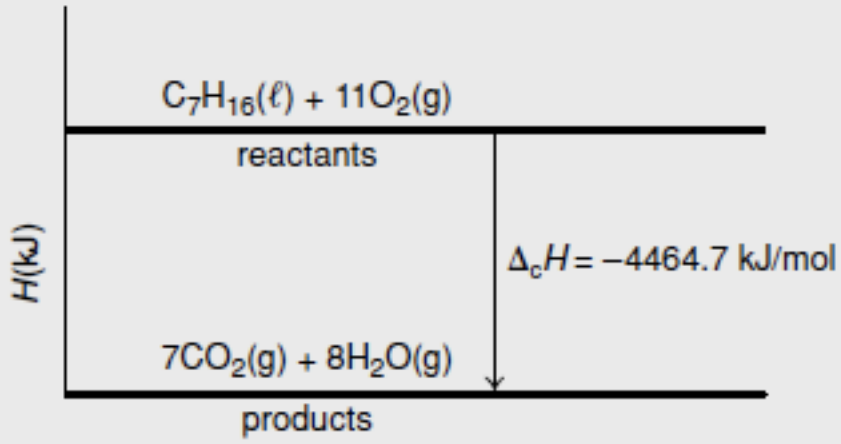
(c) $C_6H_{14}(\ell) + \frac{19}{2}O_2(g) \rightarrow 6CO_2(g) + 7H_2O(g)$

$\Delta_c H = -3854.9$ kJ/mol

(d) $C_8H_{18}(\ell) + \frac{25}{2}O_2(g) \rightarrow 8CO_2(g) + 9H_2O(g)$

$\Delta_c H = -5074.1$ kJ/mol

Q6.



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For full solutions to the practice problems, visit **www.albertachemistry.ca,** Online Learning Centre, Instructor Edition, Full Solutions.

1. –7.17 kJ
2. **(a)** –405.60 kJ/mol $O_2(g)$
   (b) –648.96 kJ/mol $CO_2(g)$
   (c) –540.8 kJ/mol $H_2O(\ell)$
3. **(a)** -2.553×10^3 kJ
   (b) -6.04×10^4 kJ
   (c) -1.044×10^6 kJ
4. –50.0 kJ
5. 1.03 kg
6. **(a)** –227 kJ/mol $NH_3(g)$
   (b) –181 kJ/mol $O_2(g)$
   (c) –227 kJ/mol NO(g)
   (d) –151 kJ/mol $H_2O(g)$

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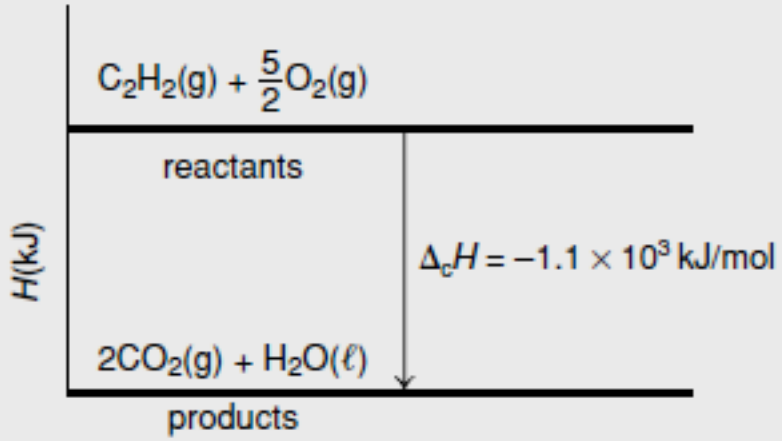
1. The first law of thermodynamics states that any change in energy of a system is accompanied by an equal and opposite change in energy of the surroundings.
   $\Delta E_{system} = -\Delta E_{surroundings}$
2. If a reaction is endothermic, more energy is consumed in breaking bonds than is released when bonds are formed.
3. **(a)** $Ag(s) + \frac{1}{2}Cl_2(g) \rightarrow AgCl(s)$ $\Delta_f H^\circ = -127.0$ kJ/mol
   (b) $C_2H_4(g) + 3O_2(g) \rightarrow 2CO_2(g) + 2H_2O(g)$
   $\Delta_c H = -1322.9$ kJ/mol
   (c) $NaCl(s) \rightarrow Na^+(aq) + Cl^-(aq)$ $\Delta H_{sol} = -44.2$ kJ/mol

4. (a) $C_2H_2(g) + \frac{5}{2}O_2(g) \rightarrow 2CO_2(g) + H_2O(\ell)$

$\Delta_c H = -1.25 \times 10^3$ kJ/mol

(b)



(c) Method 1 Using Formulas

$$n = \frac{m}{M}$$

$$= \frac{2.17 \text{ g } \cancel{C_2H_2(g)}}{\frac{26.04 \text{ g } \cancel{C_2H_2(g)}}{1 \text{ mol } \cancel{C_2H_2(g)}}}$$

$$= 0.0833 \text{ mol } C_2H_2(g)$$

$$\Delta H = n\Delta_c H$$

$$= (0.0833 \cancel{\text{ mol } C_2H_2(g)}\left(\frac{-1.25 \times 10^3 \text{ kJ}}{\cancel{\text{mol } C_2H_2(g)}}\right)$$

$$= -104 \text{ kJ}$$

Method 2 Using Dimensional Analysis

$$\text{x kJ} = 2.17 \cancel{\text{ g } C_2H_4(g)} \times \frac{1 \cancel{\text{ mol } C_2H_2(g)}}{26.04 \cancel{\text{ g } C_2H_2(g)}} \times \frac{-1.25 \times 10^3 \text{ kJ}}{\cancel{\text{mol } C_2H_2(g)}}$$

$$= -104 \text{ kJ}$$

5. In exothermic reactions, the potential energy of the products is less than the potential energy of the reactants. If the potential energy of the system decreases, the ΔH is negative and the energy of the surroundings increases. Therefore, the reaction is exothermic.

6. (a) $11H_2O\,(\ell) + 12CO_2(g) \rightarrow C_{12}H_{22}O_{11}(s) + 12O_2(g) + 2226.1\ kJ$

(b) $11H_2O\,(\ell) + 12CO_2(g) \rightarrow C_{12}H_{22}O_{11}(s) + 12O_2(g)\ \Delta_f H^\circ = -2226.1\ kJ/mol$

7. (a) $H_2S(g) + \frac{3}{2}O_2(g) \rightarrow H_2O(g) + SO_2(g)$

$\Delta_c H = -519\ kJ/mol$

(b) Method 1 Using Formulas

$$n = \frac{m}{M}$$

$$= \frac{15.0\ \cancel{g\ H_2S(g)}}{\dfrac{34.09\ \cancel{g\ H_2S(g)}}{1\ mol\ H_2S(g)}}$$

$$= 0.440\ mol\ H_2S(g)$$

$$\Delta H = n\Delta_c H$$

$$= (0.440\ \cancel{mol\ H_2S(g)})\left(\frac{-519\ kJ}{\cancel{mol\ H_2S(g)}}\right)$$

$$= -228\ kJ$$

Method 2 Using Dimensional Analysis

$$x\ kJ = 15.0\ \cancel{g\ H_2S(g)} \times \frac{1\ \cancel{mol\ H_2S(g)}}{34.09\ \cancel{g\ H_2S(g)}} \times \frac{-519\ kJ}{\cancel{mol\ H_2S(g)}}$$

$$= -228\ kJ$$

(c) Method 1 Using Formulas

$$\Delta H = n\Delta_c H$$

$$n = \frac{\Delta H}{\Delta_c H}$$

$$= \frac{-47.2 \text{ kJ}}{-519 \text{ kJ/mol } SO_2(g)}$$

$$= 0.0909 \text{ mol } SO_2(g)$$

Since 1 mol of gas at STP has a volume of 22.4 L, or molar volume is 22.4 L/mol,

$$V = nMV$$

$$= 0.0909 \text{ mol } SO_2(g) \times \frac{22.4 \text{ L } SO_2(g)}{\text{mol } SO_2(g)}$$

$$= 2.04 \text{ L } SO_2(g)$$

Method 2 Using Dimensional Analysis

$$x \text{ L } SO_2(g) = -47.2 \text{ kJ} \times \frac{1 \text{ mol } SO_2(g)}{-519 \text{ kJ}} \times \frac{22.4 \text{ L } SO_2(g)}{\text{mol } SO_2(g)}$$

$$= 2.04 \text{ L } SO_2(g)$$

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7. –49.8 kJ/mol HCl(aq)

8. (a) -4.1×10^2 kJ/mol Mg(s)

(b) All the thermal energy released by the reaction is absorbed by the solution; the solution has the same heat capacity as water; the reaction proceeds to completion.

9. -2.9×10^2 kJ/mol Na(s)

10. 29.7 °C

11. $HBr(aq) + KOH(aq) \rightarrow KBr(aq) + H_2O(\ell)$

$\Delta_r H = -57$ kJ

12. -41.9×10^2 kJ/mol NaOH(aq) or
-41.9×10^2 kJ/mol HCl(aq)

Answers to Questions for Comprehension

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Q7. Calorimetry is the use of a device called a calorimeter to study the energy changes associated with physical and chemical processes.

Q8. The law of conservation of energy states that the energy of the system plus its surroundings is constant. As energy is released (or absorbed) by a chemical reaction in a calorimeter, it results in a change in the temperature of the calorimeter surrounding it. The enthalpy change of the reaction can be calculated based on the thermal energy transfer in the calorimeter.

Q9. The second law of thermodynamics states that thermal energy is spontaneously transferred from an object at a higher temperature to an object at a lower temperature until the two objects reach the same temperature. Therefore, as energy is released (or absorbed) by a reaction in a calorimeter, heat will be transferred to (or from) the calorimeter and the enthalpy change can be measured.

Q10. A calorimeter will allow you to measure the enthalpy change of a reaction if the system is isolated, the amount of energy exchanged with the calorimeter is either measurable or negligible, any solutions involved in the reaction can be assumed to have the same properties as water, and finally, the reaction occurs at a constant pressure.

Section 9.2 Review Answers

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1. A calorimeter should isolate the system from the surroundings (not allow heat or matter transfer) and it should either have a known specific heat capacity, or absorb so little energy that it is negligible.

2. Method 1 Using Formulas

$$Q = C\Delta t$$
$$= \left(\frac{7.61\ \text{kJ}}{^\circ\text{C}}\right)(5.23\ ^\circ\text{C})$$
$$= 39.8\ \text{kJ}$$

Method 2 Using Dimensional Analysis

$$\text{x kJ} = \frac{7.61\ \text{kJ}}{^\circ\text{C}} \times 5.23\ ^\circ\text{C}$$
$$= 39.8\ \text{kJ}$$

The enthalpy change of the reaction is $\Delta H = -39.8$ kJ.

3. $Q = mc\Delta t$

$$= (150\ \text{g})\left(\frac{4.19\ \text{J}}{\text{g}^\bullet\ ^\circ\text{C}}\right)(-5.0\ ^\circ\text{C})$$
$$= -3.1\ \text{kJ}$$

4. A polystyrene cup is suitable for a constant pressure calorimeter because it is an excellent insulator. Polystyrene tends not to react with most chemicals used in the high school laboratory. Polystyrene is hydrophobic, so the cup can usually be dried with vigorous shaking.

5. If you use a concentrated reactant solution in a coffee cup calorimeter, you cannot make the same assumptions as if you used a dilute reactant. A dilute aqueous solution is assumed to have the same density and specific heat capacity as pure water. These assumptions do not hold true for concentrated aqueous solutions, as the densities and specific heat capacities would be significantly different from the density and specific heat capacity of water.

6. **(a)** Equipment and chemicals should include safety equipment, materials to build a coffee cup calorimeter (a cup, lid, thermometer and stirrer), sulfuric acid, and water.

   (b) Students' procedures should demonstrate an understanding of calorimetry and concern for safety. Ensure students are adding acid to water, not the other way around. The procedure should involve first finding the temperature of water that has equilibrated to room temperature. Then students should suggest adding a small, measured volume of concentrated sulfuric acid at room temperature to a measured volume of water at room temperature in a calorimeter, and then monitoring the temperature change. Students should point out that concentrated sulfuric acid is extremely corrosive and dangerous.

(c)

	Volume of solution (mL)	Initial temperature (°C)	Final temperature (°C)
sulfuric acid			
water			

(d) Students need to know the density of concentrated sulfuric acid and the concentration of concentrated sulfuric acid. Students may add that they need to know the specific heat capacity of the sulfuric acid solution. If not, then they must include the simplifying assumption that the specific heat capacity of the solution is the same as that of water.

(e) Students assume that energy is conserved and that no heat is absorbed by the calorimeter or the surroundings. They also assume that the specific heat capacity of the solution is the same as the specific heat capacity of water (unless it is in the list of needs above).

7. (a) The reactants are in stoichiometric quantities, so a calculation for limiting reagent is not required.

$$Q = mc\Delta t$$

$$= (200\ \cancel{g})\left(\frac{4.19\ \text{J}}{\cancel{g}^{\bullet}\ ^{\circ}\cancel{\text{C}}}\right)(21.34\ ^{\circ}\cancel{\text{C}} - 21.01\ ^{\circ}\cancel{\text{C}})$$

$$= 276\ \text{J}$$

The thermal energy of the reaction is therefore –276 J.

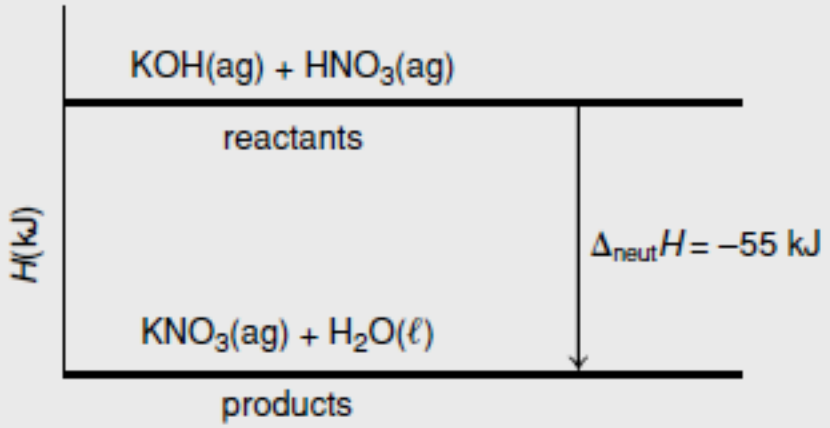
$$\Delta H = n\Delta_r H$$

$$\Delta_r H = \left(\frac{\Delta H}{n}\right)$$

$$= \frac{(-276\ \text{J})}{\left(\frac{0.050\ \text{mol KOH(aq)}}{\cancel{\text{L}}}\right)(0.100\ \cancel{\text{L}})}$$

$$= -55\ \text{kJ/mol KOH(aq) or } HNO_3\text{(aq)}$$

(b) $KOH(aq) + HNO_3(aq) \rightarrow KNO_3(aq) + H_2O(\ell)$
$\Delta_{neut}H = -55$ kJ/mol



(c) The students should indicate that they would repeat the experiment a few times to reduce random error. If the experiment was started just below room temperature, then the final temperature would be slightly above. This should minimize error due to heat loss/gain to and from the surroundings as both conditions should cancel each other out.

8. Students should suggest that the human subject be kept in a room that is insulated from its surroundings. Students may suggest that the room should include exercise equipment, a bed, and adequate food and water. The room would need to be installed with a very sensitive thermometer to measure temperature changes to 1/100 °C. Students will probably suggest that the subject be placed in this room for a period of time and the energy given up by his/her body to the circulating air be monitored over time. Accept all thoughtful proposals that show an understanding of how calorimetry works in general.

Chapter 9 Review Answers

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Answers to Understanding Concepts Questions

1. Any change of energy in the system is accompanied by an equal and opposite energy change in the surroundings.
$\Delta E_{system} = -\Delta E_{surroundings}$

2. (a) Heat capacity, C, is the quantity of energy required to raise the temperature of a substance or object by 1 °C, whereas specific heat capacity, c, is the quantity of energy to raise 1 g of a pure substance by 1 °C. Heat capacity is generally used for a bomb calorimeter or a calorimeter that has been calibrated, whereas specific heat capacity is usually used for substances that are part of the system, often water.

(b) Enthalpy change, ΔH, is the potential energy change of a system during a process, molar enthalpy change is the enthalpy change of the system expressed per mole of one of the substances involved. Enthalpy change is most often for an equation balanced with the smallest whole numbers.

enthalpy change: $C_6H_{12}O_6(s) + 6O_2(g) \rightarrow 6CO_2(g) + 6H_2O(\ell)$ $\Delta H = -2870$ kJ

molar enthalpy change: $C_6H_{12}O_6(s) + 6O_2(g) \rightarrow 6CO_2(g) + 6H_2O(\ell)$ $\Delta H = -2870$ kJ/mol $C_6H_{12}O_6(s)$ or -478.8 kJ/mol $O_2(g)$

(c) Formation reactions describe the production of one mole of a substance from its elements in standard state, whereas a combustion reaction is the complete combustion of one mole of a substance in the presence of oxygen. For example, in methane's case,

formation: $C(s) + 2H_2(g) \rightarrow CH_4(g)$

combustion: $CH_4(g) + 2O_2(g) \rightarrow CO_2(g) + 2H_2O(g)$ (the water could be liquid or vapour form, depending on the conditions)

3. Kinetic energy is energy in motion. It is measured as a temperature change. Potential energy is stored in the intermolecular or intramolecular bonds of substances. It is measured as a change in temperature of the surroundings when bonds are broken or formed.

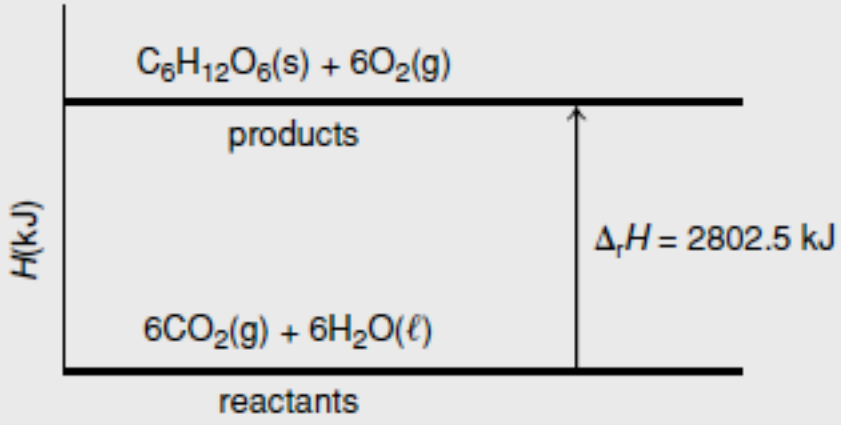
(a) heating water

(b) boiling water

(c) heating water from 25 °C to steam at 100 °C

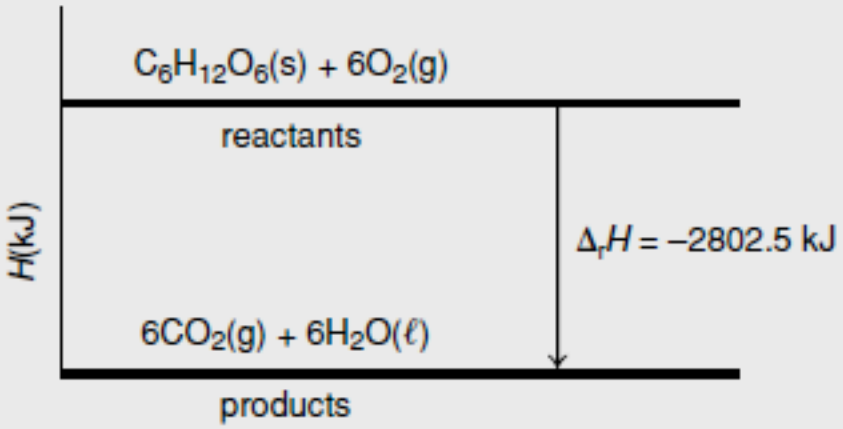
4. The nickel will show a larger increase in temperature because the specific heat capacity is lower.

5. (a) $6CO_2(g) + 6H_2O(\ell) \rightarrow C_6H_{12}O_6(s) + 6O_2(g)$
$\Delta H = 2802.5$ kJ



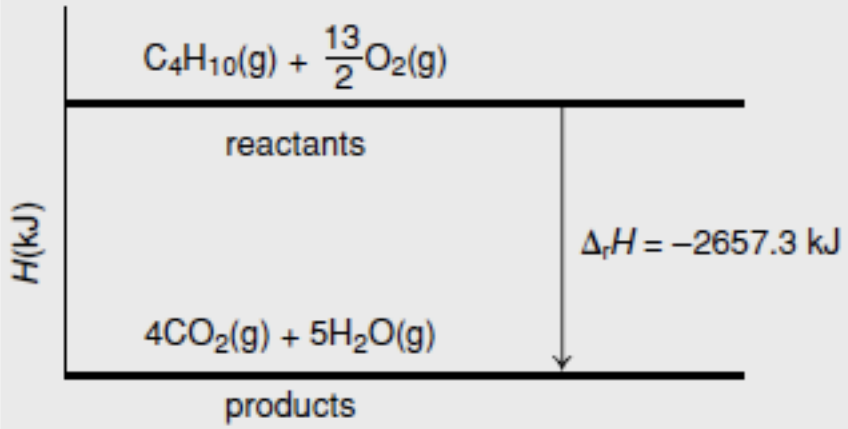
(b) $C_6H_{12}O_6(s) + 6O_2(g) \rightarrow 6CO_2(g) + 6H_2O(\ell)$

$\Delta H = -2802.5$ kJ



(c) $C_4H_{10}(g) + \frac{13}{2}O_2(g) \rightarrow 4CO_2(g) + 5H_2O(g)$

$\Delta H = -2657.3$ kJ



6. (a) $C_7H_{16}(\ell) + 11O_2(g) \rightarrow 7CO_2(g) + 8H_2O(g)$

$\Delta_c H = -4464.7$ kJ/mol $C_7H_{16}(\ell)$

(in an open system – gaseous water produced)

$C_7H_{16}(\ell) + 11O2(g) \rightarrow 7CO_2(g) + 8H_2O(\ell)$

$\Delta_c H = -4816.7$ kJ/mol $C_7H_{16}(\ell)$

(in a closed system – liquid water produced)

(b) $2C_4H_{10}(g) + 13O_2(g) \rightarrow 8CO_2(g) + 10H_2O(g)$

$\Delta_c H = -2657.3$ kJ/mol $C_4H_{10}(g)$

(in an open system – gaseous water produced)

$2C_4H_{10}(g) + 13O_2(g) \rightarrow 8CO_2(g) + 10H_2O(\ell)$

$\Delta_c H = -2877.3$ kJ/mol $C_4H_{10}(g)$

(in a closed system – liquid water produced)

7. Polystyrene cups are suitable for a constant pressure calorimeter because polystyrene is an excellent insulator. When nested, the air pocket in-between further insulates the calorimeter from the surroundings. Polystyrene tends not to react with most chemicals used in the high school laboratory. Polystyrene is hydrophobic, so the cup can usually be dried with vigorous shaking.

8. Experiments using a coffee cup calorimeter assume it is a perfectly insulated container, so there are no heat gains or losses. Secondly, it is assumed that the heat capacity of the calorimeter, which includes the thermometer and stirrer, can be neglected. Heat will be transferred to or from water or an aqueous solution in the calorimeter, and it is assumed that the heat capacity of the fluid is the same as that of pure water. If the volume of an aqueous solution is measured, the assumption is made that the density of the solution is identical with that of pure water so that the mass of solution can be calculated.

9. Juice cans do not make very good calorimeters because they do not effectively isolate the system from the surroundings. Polystyrene cups, however, cannot be used for investigations involving combustion, so juice cans are a simple, but not perfect, substitute.

10. **(a)** exothermic

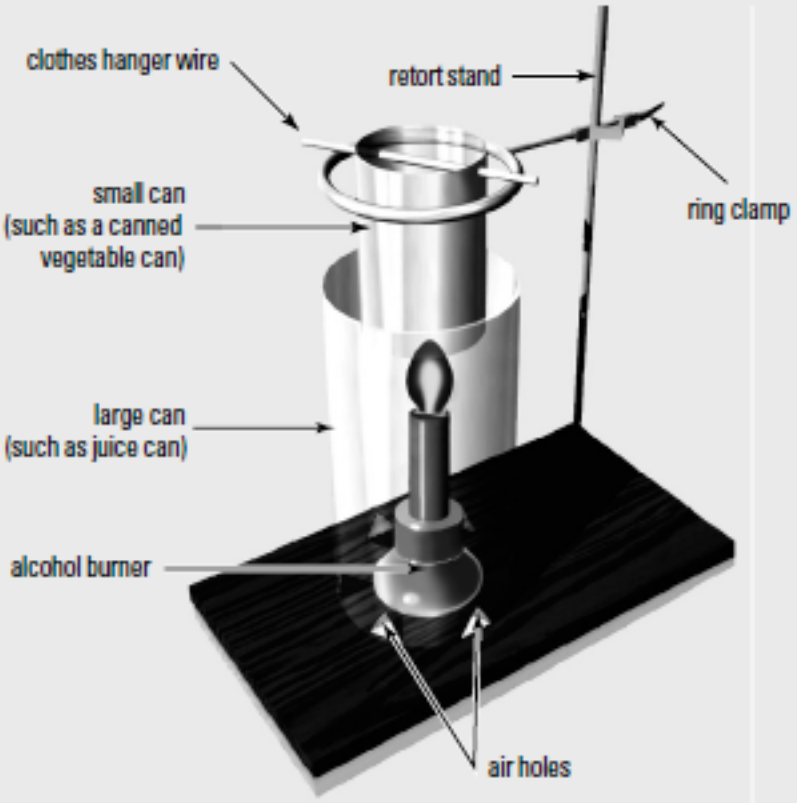
(b) endothermic

(c) exothermic

Answers to Applying Concepts Questions

11. Water is a more suitable solvent in a simple calorimeter because it will dissolve more solutes that are commonly used in high school chemistry labs and it has a high specific heat capacity.

12. (a)



(b), (c) Accept all reasonable procedures. Students will probably design an experiment similar to Investigation 9.B, using a measured mass of ethanol in the lamp to heat water in a calorimeter. They can determine the enthalpy change based on the temperature change of the water.

(d) Students should state that they would assume there are no heat losses to the surroundings, and the heat content of the calorimeter is negligible (if not, they must include the specific heat capacity of the calorimeter in the equation below). They will need to assume that the fuel is pure ethanol. They will also need to assume that complete combustion takes place, although this is probably not the case.

(e) mass of $C_2H_5OH(\ell)$ burned = mass of burner before – mass of burner after

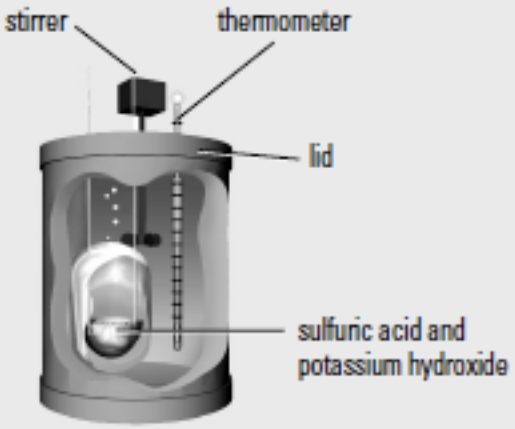
moles (n) of $C_2H_5OH(\ell)$ = mass of $\frac{C_2H_5OH(\ell)(g)}{M}$,

where M = molar mass of $C_2H_5OH(\ell)$ = 46.08 g/mol

$Q = mc\Delta t$ (may also have another $mc\Delta t$ term for the calorimeter), where m = mass of water in the calorimeter, c = 4.19 J/g•°C, Δt = change in temperature of the water (°C)

$\Delta H = n\Delta_r H$ where n = the moles of $C_2H_5OH(\ell)$, ΔH is the thermal energy change of the system which is –(thermal energy change of the surroundings).

13. (a)



(b), (c) Accept all reasonable procedures. Students will probably design an experiment similar to Investigation 9.A, using two solutions. Since sulfuric acid, $H_2SO_4(aq)$, provides two moles of hydrogen ions to the one mole of hydroxide ions of potassium hydroxide, KOH(aq), the volume (or concentration) of the potassium hydroxide should be double that of sulfuric acid to be in stoichiometric quantities. Students can determine the enthalpy change based on the temperature change of the solutions.

(d) Students should state that they would assume there are no heat losses to the surroundings, and the heat content of the calorimeter is negligible. They will need to assume the density and specific heat capacity of the solutions is the same as that of water.

(e) If students do not use twice as much (or double the concentration) of the potassium hydroxide solution, they must do a limiting reagent calculation to determine the number of moles of acid or base that react. They should use the equations:

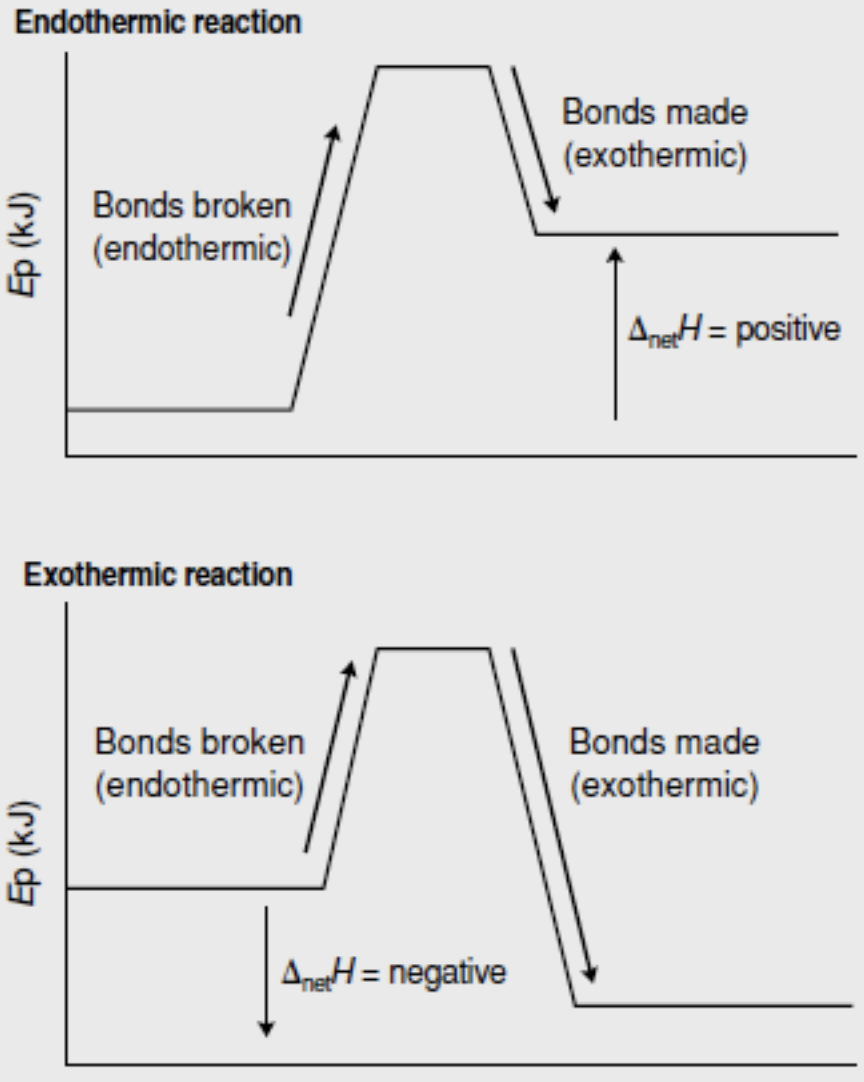
$Q = mc\Delta t$ and $\Delta H = n\Delta_r H$, where

n is the moles of hydrogen or hydroxide ions, m is the total mass of both solutions (the mass of water), c is the specific heat capacity of water, 4.19 J/g•°C, Δt = change in temperature of the two solutions (°C).

14. As the number of carbons in a hydrocarbon increases, the molar enthalpy of combustion increases. In a homologous series, that increase will be proportional to the number of carbons in the hydrocarbon.

15. All chemical reactions involve bonds being broken and other bonds being made. Breaking bonds is endothermic, making bonds is exothermic. The sum of the two energy changes determines whether the overall reaction will be endothermic or exothermic. If more energy is required to break the bonds than is released when making new bonds, then the overall reaction is endothermic. If more energy is released when making bonds than is required to break the bonds, then the overall reaction will be exothermic. These concepts are summarized in the diagram below:

***Ep* diagrams summarizing endo and exothermic reactions in steps**



Answers to Solving Problems Questions

16. $Q = mc\Delta t$

$$= (1.00\text{ kg})\left(\frac{4.19\text{ J}}{\text{g}^\bullet\,^\circ\text{C}}\right)(99.0\,^\circ\text{C} - 22.0\,^\circ\text{C})\left(\frac{1000\text{ g}}{1\text{ kg}}\right) \times \left(\frac{1\text{ kJ}}{1000\text{ J}}\right)$$

$$= 322\text{ kJ}$$

17. Method 1 Using Formulas

$$n = \frac{m}{M}$$

$$= \frac{10.0\text{ g } CH_3COOH(\ell)}{\frac{60.06\text{ g } CH_3COOH(\ell)}{\text{mol } CH_3COOH(\ell)}}$$

$$= 0.1665\text{ mol } CH_3COOH(\ell)$$

$$\Delta H = n\Delta_c H$$

$$\Delta_c H = \frac{\Delta H}{n}$$

$$= \frac{-144.77\text{ kJ}}{0.1665\text{ mol } CH_3COOH(\ell)}$$

$$= -869\text{ kJ/mol } CH_3COOH(\ell)$$

Method 2 Using Dimensional Analysis

$$x\,\frac{\text{kJ}}{\text{mol } CH_3COOH(\ell)} = -144.77\text{ kJ} \times$$

$$\frac{60.06\text{ g } \cancel{CH_3COOH(\ell)}}{\text{mol } CH_3COOH(\ell)} \times \frac{1}{10.0\text{ g } \cancel{CH_3COOH(\ell)}}$$

$$= -869\text{ kJ/mol } CH_3COOH(\ell)$$

18. (a) $CH_3OH(\ell) + \frac{3}{2}O_2(g) \rightarrow$

$$CO_2(g) + 2H_2O(g) \quad \Delta_c H = -727 \text{ kJ/mol}$$

(b) Method 1 Using Formulas

$$n = \frac{m}{M}$$

$$= \frac{44.3 \text{ g } \cancel{CH_3OH(\ell)}}{\dfrac{32.05 \text{ g } \cancel{CH_3OH(\ell)}}{\text{mol } CH_3OH(\ell)}}$$

$$= 1.38 \text{ mol } CH_3OH(\ell)$$

$$\Delta H = n\Delta_c H$$

$$= (1.38 \text{ } \cancel{\text{mol } CH_3OH(\ell)}) \left(\frac{-727 \text{ kJ}}{\cancel{\text{mol } CH_3OH(\ell)}} \right)$$

$$= -1.00 \times 10^3 \text{ kJ}$$

Method 2 Using Dimensional Analysis

$$\text{x kJ} = \frac{-727 \text{ kJ}}{\cancel{\text{mol } CH_3OH(\ell)}} \times 44.3 \text{ } \cancel{\text{g } CH_3OH(\ell)} \times \frac{1 \text{ } \cancel{\text{mol } CH_3OH(\ell)}}{32.05 \text{ } \cancel{\text{g } CH_3OH(\ell)}}$$

$$= -1.00 \times 10^3 \text{ kJ}$$

19. (a) 34.2 g $CaC_2(s)$

(b) 599 kJ

(c) 1.14×10^6 kJ

20. mass of wax burned = 17.36 g – 17.01 g = 0.35 g

$$n = \frac{m}{M}$$

$$= \frac{0.35 \text{ } \cancel{\text{g } C_{25}H_{52}(s)}}{\dfrac{352.77 \text{ } \cancel{\text{g } C_{25}H_{52}(s)}}{\text{mol } C_{25}H_{52}(s)}}$$

$$= 9.9 \times 10^{-4} \text{ mol } C_{25}H_{52}(s)$$

$$Q = mc\Delta t_{water} + mc\Delta t_{aluminium\ can}$$

$$= (125.03\ \cancel{g})\left(\frac{4.19\ J}{\cancel{g}^{\bullet}\cancel{^{\circ}C}}\right)(35.7\ \cancel{^{\circ}C} - 14.3\ \cancel{^{\circ}C}) +$$

$$(35.03\ \cancel{g})\left(\frac{0.897\ J}{\cancel{g}^{\bullet}\cancel{^{\circ}C}}\right)(35.7\ \cancel{^{\circ}C} - 14.3\ \cancel{^{\circ}C})$$

$$= 11\ 211\ J + 672\ J$$

$$= 11.9\ kJ$$

Therefore, $\Delta H = -11.9$ kJ.

$$\Delta H = n\Delta_c H$$

$$\Delta_c H = \frac{\Delta H}{n}$$

$$= \frac{-11.9\ kJ}{9.9 \times 10^{-4}\ mol\ C_{25}H_{52}(s)}$$

$$= -1.2 \times 10^4\ kJ/mol\ C_{25}H_{52}(s)$$

The accepted value for the molar enthalpy of combustion for paraffin is about -1.8×10^4 kJ/mol. The experimental value is lower than expected. This could be due to heat lost to the surroundings (which would decrease the value of the numerator), or incorrectly massing the water, metal can or candle. Perhaps the can was not made of pure aluminium, which would change the value of the specific heat capacity for the can.

$$\text{Percentage Difference} = \frac{\text{Accepted} - \text{Experimental}}{\text{Accepted}} \times 100\%$$

$$= \frac{-1.8 \times 10^4\ \cancel{kJ/mol} - (-1.2 \times 10^4\ \cancel{kJ/mol})}{-1.8 \times 10^4\ \cancel{kJ/mol}} \times 100\%$$

$$= 33\%$$

21. $$Q = mc\Delta t$$

$$= (500\ g)\left(\frac{4.19\ \cancel{J}}{\cancel{g}^{\bullet}\ \cancel{^{\circ}C}}\right)(13.5\ ^{\circ}C)\left(\frac{1\ kJ}{1000\ \cancel{J}}\right)$$

$$= 28.3\ kJ$$

Therefore, change in enthalpy of reaction is $\Delta H = -28.3$ kJ.

$$\Delta H = n\Delta_c H$$

$$n = \frac{\Delta H}{\Delta_c H}$$

$$= \frac{-28.3 \cancel{\text{kJ}}}{-3508.8 \cancel{\text{kJ}}/\text{mol } C_5H_{12}(\ell)}$$

$$= 8.07 \times 10^{-3} \text{ mol } C_5H_{12}(\ell)$$

$$n = \frac{m}{M}$$

$$m = nM$$

$$= (8.07 \times 10^{-3} \cancel{\text{mol } C_5H_{12}(\ell)})\left(\frac{72.17 \text{ g } C_5H_{12}(\ell)}{\cancel{\text{mol } C_5H_{12}(\ell)}}\right)$$

$$= 0.582 \text{ g } C_5H_{12}(\ell)$$

22.

$$n = \frac{m}{M}$$

$$= \frac{0.930 \cancel{\text{ g C(s)}}}{\frac{12.01 \cancel{\text{ g C(s)}}}{\text{mol C(s)}}}$$

$$= 0.0774 \text{ mol C(s)}$$

$$\Delta H = n\Delta_c H$$

$$= (0.0774 \text{ mol C(s)})(-394 \text{ kJ/mol C(s)})$$

$$= -30.5 \text{ kJ}$$

The thermal energy change in the calorimeter is 30.5 kJ. Since a bomb calorimeter's heat capacity is not a specific heat capacity, its heat capacity is expressed in kJ/°C, not kJ/g•°C, so the formula should be changed to $Q = c\Delta t$ instead of $Q = mc\Delta t$.

$$Q = c\Delta t$$

$$c = \frac{Q}{\Delta t}$$

$$= \frac{(30.5\ \text{kJ})}{(28.15\ ^\circ\text{C} - 25.00\ ^\circ\text{C})}$$

$$= 9.68\ \text{kJ/}^\circ\text{C}$$

The increase in temperature is evidence that the reaction is exothermic.

Answers to Making Connections Questions

23. a) $$\text{x minutes} = 527.18\ \text{kJ} \times \frac{1\ \text{minute}}{20\ \text{kJ}}$$

$$= 26\ \text{minutes}$$

b) $$\text{x g} = \frac{37.5\ \text{g}}{527.18\ \cancel{\text{kJ}}} \times \frac{20\ \cancel{\text{kJ}}}{1\ \cancel{\text{minute}}} \times \frac{60\ \cancel{\text{minutes}}}{1\ \text{h}} \times 4.0\ \text{h}$$

$$= 3.4 \times 10^2\ \text{g}$$

c) $$\text{x apples} = \frac{1\ \text{apple}}{283\ \cancel{\text{kJ}}} \times \frac{20\ \cancel{\text{kJ}}}{1\ \cancel{\text{minute}}} \times \frac{60\ \cancel{\text{minutes}}}{1\ \text{h}} \times 4.0\ \text{h}$$

$$= 17\ \text{apples}$$

24. (a) $H_2(g) + O_2(g) \rightarrow H_2O(g)\ \Delta H = -241.8\ \text{kJ}$

(b) $8.3 \times 10^4\ \text{L}$

(c) For hydrogen gas combustion, there are no carbon dioxide or other greenhouse gas emissions; the only by-product is pure water; and it is extremely efficient.

(d) Hydrogen gas requires large storage space and a large engine.